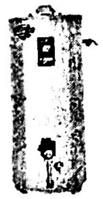
#1

Thermochemistry Problem Set

Problem 1

An average hot water tank will hold 300.0 L of water needs to be kept at 65.0°C in order to kill any bacteria in the water. If you are filling a new tank with water at a starting temperature of 12.5°C, what mass of natural gas (methane) must be burned in a complete combustion reaction in order bring your tank up to the correct temperature?



$$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}$$

① Find ΔH°_r using enthalpy of formation

$$\Delta H^\circ_r = (\Delta H^\circ_f CO_{2(g)} + 2\Delta H^\circ_f H_2O_{(g)}) - (\Delta H^\circ_f (CH_4)_g)$$

$$= [-393.5 + 2(-241.8) - (-74.6)]$$

$$= -802.5 \text{ kJ}$$

② $Q_{water} = 300\,000 \text{ g} \times 4.184 \text{ J/g°C} \times 52.5 \text{ °C}$

65.0
−12.5

300.0 L × 1.0 g/mL × 1000 mL/L

$= 65\,898\,000 \text{ J}$ or $65\,898 \text{ kJ}$

③ $\Delta H = -Q_{sur} = -65\,898 \text{ kJ}$

④ $\frac{\Delta H_1}{n_1} = \frac{\Delta H_2}{n_2} \Rightarrow \frac{-65898 \text{ kJ}}{x \text{ mol}} = \frac{-802.5 \text{ kJ}}{1 \text{ mol}}$

$x = \frac{-65898}{-802.5}$

$x = 82.116 \text{ mol}$

⑤ $\text{mass} = n \times M$

$= 82.116 \text{ mol} \times 16.04 \text{ g/mol}$

$= 1317.1 \text{ g}$

Correct to 3 sig fig ⇒ 1320 g of CH_4 is needed

Problem 2

The Thermite reaction can be used to produce molten iron for welding railway tracks together.

$$Fe_2O_{3(s)} + 2\,Al_{(s)} \rightarrow Al_2O_{3(s)} + 2\,Fe_{(s)}$$

Calculate the amount of heat produced using 45.0 g of aluminum and excess iron (III) oxide.

① $\Delta H_r^\circ = [\Delta H_f^\circ Al_2O_{3(s)}] - [\Delta H_f^\circ Fe_2O_{3(s)}]$

$= -1675.7 - (-824.2)$

$= -851.5\ kJ$

② $mol\ Al = 45.0g \times \frac{1\ mol}{26.98g}$

$= 1.6679\ mol$

③ $\frac{\Delta H_1}{n_1} = \frac{\Delta H_2}{n_2}$

$\frac{-851.5\ kJ}{2\ mol\ Al} = \frac{x}{1.6679\ mol}$

Mole Coefficient in balanced equation

$x = \frac{-851.5 \times 1.6679}{2}$

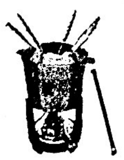
$= -710.1\ kJ$

3 sig fig ⇒ $-7.10 \times 10^2\ kJ$

#3

Problem 3

It's 1976 and you are throwing a fondue party! You need to heat 800.0 mL of vegetable oil to 375°F (190.0°C) using an ethanol burner. Assuming a perfect transfer of heat from the burner into the oil, what mass of ethanol is required? Assume your kitchen is at SATP, vegetable oil has a density of 0.920 g/mL and the specific heat capacity of vegetable oil is 2.000 J/g•°C



$C_2H_5OH_{(l)} + 3O_{2(g)} \rightarrow 2CO_{2(g)} + 3H_2O_{(g)}$

① $Q_{oil} = mc\Delta T$
$= 736 \times 2.000 \times 165$
$= 242\,880\,J$

$m_{oil} = 800.0\,mL \times 0.920\,g/mL$
$= 736\,g$

$\Delta T = 190.0° - 25°C$
$= 165°C$ ↑SATP

② $\Delta H_{system} = -242\,880\,J$ or $-242.880\,kJ$

③ find $\Delta H_r°$ using $\Delta H_f°$

$\Delta H_r° = \left(2\Delta H_f°\,CO_{2(g)} + 3\Delta H_f°\,H_2O_{(g)}\right) - \left[\Delta H_f°\,C_2H_5OH_{(l)}\right]$

$= \left[2(-393.5) + 3(-241.8)\right] - \left[(-277.6)\right]$

$= -1234.8\,kJ$

④ $\frac{\Delta H_1}{n_1} = \frac{\Delta H_2}{n_2}$

$\frac{-1234.8\,kJ}{1\,mol} = \frac{-242.880\,kJ}{n_2}$

Found in steps 1 + 2

Mole coefficient in Balanced equation for ethanol

$n_2 = 0.197\,mol$

⑤ $m = n \times M \Rightarrow n = 0.197\,mol \times 46.08\,g/mol$
$= 9.08\,g$

2 sig dig = 9.1 g

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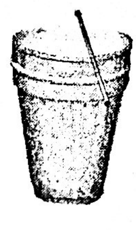
~~Problem 4~~

100.0 mL of 3.0 M HCl can be neutralized with exactly 200.0 mL of 1.5 M NaOH. If this reaction is done in a coffee cup calorimeter, what temperature change would you expect to observe?



$$NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

1 $\Delta H^\circ_r = \left[\Delta H^\circ_f\, NaCl_{(aq)} + \Delta H^\circ_f\, H_2O_{(l)}\right] - \left[\Delta H^\circ_f\, NaOH_{(aq)} + \Delta H^\circ_f\, HCl\right]$

$= (-411.2 + (-285.8)) - ((-92.3) + (-425.6))$

$= -179.1\ kJ$

2 $\dfrac{\Delta H_1}{n_1} = \dfrac{\Delta H_2}{n_2}$

Note: $mol_{NaOH} = mol_{HCl}$ both reacting completely

$\dfrac{-179.1\ kJ}{1\ mol\ NaOH} = \dfrac{x\ kJ}{0.30\ mol}$

$x = -53.73\ kJ$

3) $Q_{surr} = +53.73\ kJ$ or $53730\ J$

4 $Q = mc\Delta T$

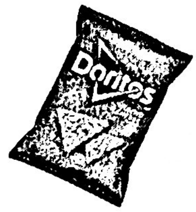
$53730\ J = 300.0\ g \times 4.184\ \frac{J}{g°C} \times \Delta T$

$\Delta T = 42.8\ °C$

2 sig. fig $\Rightarrow$ 43°C

Problem 5

A lab technician places 5.00 g of Doritos (2 chips) into a bomb calorimeter with a heat capacity of 9.23 kJ/°C. The initial temperature of the calorimeter system is 21.0°C. After burning the food, the final temperature of the system is 32.0°C. How much thermal energy is released by the combustion of the food in kilojoules per gram?



Follow-up question!
Calories on food packages are actually kilocalories (kcal). If 1 kj = 0.2388 kcal, what is the caloric value of a 45 g bag?

$Q = mc\Delta T$ or $Q = C \times \Delta T$

they gave us $9.23\ kJ/°C$

$Q = \frac{9.23\ kJ}{°C} \times 11.0°C$

$= 101.53\ kJ$ $\quad \Delta H_{syst} = -101.53\ kJ$

$\frac{kJ}{g} \Rightarrow \frac{-101.53\ kJ}{5\ g}$

$= 20.3\ kJ/g$

$\frac{20.3\ \cancel{kJ}}{g} \times \frac{0.2388\ kcal}{1\ \cancel{kJ}}$

$= 4.85\ kcal/g$

Food sample $= 45\ g$ $\quad \therefore \frac{4.85\ kcal}{\cancel{g}} \times 45\ \cancel{g}$

$= 218\ kcal$